Reaction Rates

Measuring the Rate of a Reaction

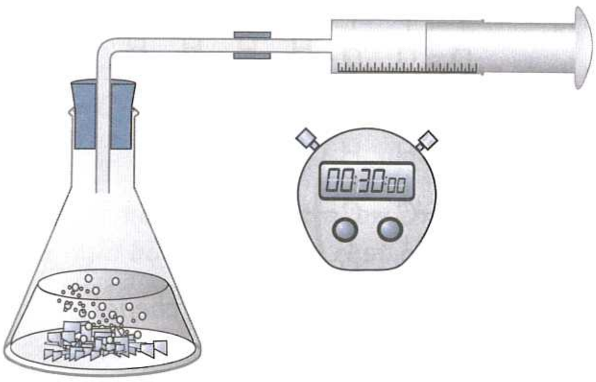
The rate of reaction is just a measure of how fast a particular reaction is going.

You need to know some of the ways that you can follow the rates of different reactions.

They're all about measuring how fast the reactants are being used up, or measuring how fast the products of the reaction are forming.

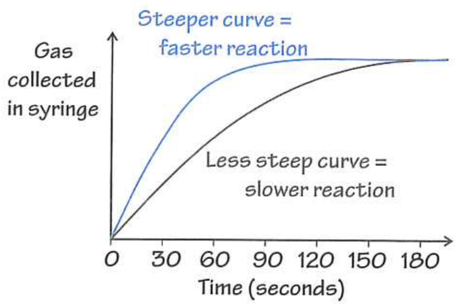
Example: Measuring the rate of reaction between hydrochloric acid and magnesium metal.

magnesium + hydrochloric acid → magnesium chloride + hydrogen



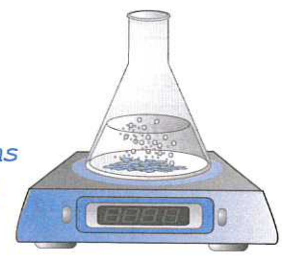
- *Use a syringe to collect the hydrogen gas that is given off during the reaction.*
- *Use a stopwatch to time the reaction.*
- *At timed intervals, say every 30 seconds, record how much hydrogen gas has been produced.*

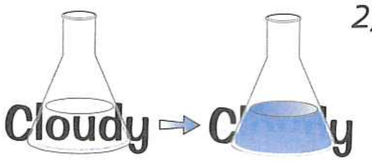
Plotting graphs lets you compare rates of reactions.



There are lots of other ways of measuring the rate of a reaction:

1) *You can measure the change in mass that occurs during a reaction where gas is released as one of the products.*
When solid calcium carbonate reacts with hydrochloric acid, carbon dioxide gas is released. If you mix the reactants in a flask and place the flask on a mass balance, you can record the decrease in mass as carbon dioxide is released.





2) *You can follow the colour change of a reaction. This includes precipitation reactions, where the solution turns cloudy as more of the product is made.*
If you mix sodium thiosulfate and hydrochloric acid, a precipitate of sulfur forms. Put a solid black mark on a piece of paper underneath the flask. Time how long it takes for the mark to be completely obscured.

3) *You can measure changes in temperature or pH that occur during the reaction.*
The reaction between sodium thiosulfate and hydrochloric acid is a neutralisation reaction. So you could also measure its rate by using a pH meter to track changes in pH.

Reaction Rates and Catalysts

Changing the Rate of Reaction

The rate of reaction depends on how often two particles collide with each other — the more collisions that occur between particles the faster the rate of reaction. But the particles also need to collide in the right direction and with enough energy or they won't react.

These factors all increase the rate of reaction:

1) Increasing temperature — the particles tend to have more kinetic energy. This means that they move around faster, and so are more likely to collide with each other and have enough energy to react.
2) Increasing concentration (or pressure in gases) — this means that the particles of reactant will be closer together, so they will be more likely to collide.
3) Increasing the surface area of a solid reactant — this increases the number of particles of the solid reactant able to come into contact with other reactants.

Catalysts Speed Up Reactions

Activation energy is the minimum amount of energy needed for a reaction to happen. This initial input of energy is needed to break bonds and start the reaction off.

A catalyst increases the rate of a reaction by lowering its activation energy. So catalysts can speed up reactions.

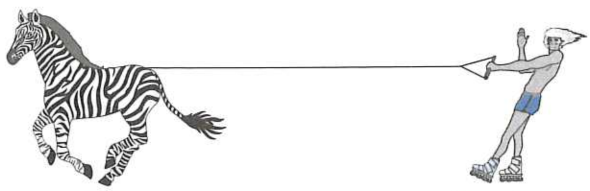
A catalyst is any substance which changes the rate of a reaction, without being changed or used up itself.

Catalysts are also very specific — most will only catalyse a single reaction.

There are loads of advantages to using catalysts:

1) Catalysts reduce the need for high temperatures and pressures in industrial reactions, like hydrocarbon cracking (see page 40) and ethanol production (see page 26). This makes these processes cheaper to run.
2) Using lower temperatures also means less energy demand, and so lower CO_2 emissions.

Living things produce enzymes which act as biological catalysts. Enzymes are protein molecules that speed up useful chemical reactions in the body.



Reversible Reactions

Reversible Reactions Go Both Ways

In a reversible reaction, the products can react with each other and change back into the reactants.

Reactants		Products
A + B	$\rightleftharpoons$	C + D

So there are actually two reactions happening at once: $A + B \rightarrow C + D$ and $C + D \rightarrow A + B$.

Example: The industrial production of ethanol from ethene.

$$H_2C{=}CH_2(g) + H_2O(g) \underset{\text{endothermic}}{\overset{\text{exothermic}}{\rightleftharpoons}} CH_3CH_2OH(g)$$

Catalyst: H_3PO_4
Temperature: 300 °C
Pressure: 60 atm

Because the reaction is reversible you don't get a high yield — some of the ethanol converts back to ethene and water. But you can keep removing and recycling any ethene that you have left, so you can still end up with a good overall yield.

Reversible Reactions Reach an Equilibrium

If a reversible reaction is taking place in a closed system it will eventually reach a state of equilibrium.

A closed system is one where none of the reactants or products can escape.

1) When a reaction begins there will be a high concentration of reactants, and a low concentration of products in the system. So the forward reaction will be fast, and the reverse reaction quite slow.
2) The concentration of reactants will gradually decrease, while the products build up. So the forward reaction will start to slow down while the reverse reaction speeds up.
3) After a while the forward reaction and the reverse reaction end up going at the same rate. From this point on the concentration of the reactants and products won't change.
4) This is called dynamic equilibrium. The forward and reverse reactions are both still happening — some reactant is being made into product, and some product is being made into reactant.
5) But since these processes are going at exactly the same rate, it seems as if nothing's happening.

Influence of Conditions on Yield

Position of Equilibrium

The position of equilibrium tells you the amount of reactants compared to the amount of products that are present when the reaction reaches an equilibrium.

$$\overset{\text{Reactants}}{A + B} \rightleftharpoons \overset{\text{Products}}{C + D}$$

If the position of equilibrium lies on the left-hand side, there are more reactants than products in the reaction mixture.

If the position of equilibrium lies on the right-hand side, there are more products than reactants in the reaction mixture.

Changing Conditions Changes the Equilibrium Position

Altering the conditions of a reversible reaction can move the position of equilibrium in one direction or the other. Careful control of the conditions can result in a higher yield (more of the products).

Look at the production of ethanol from ethene again as an example:

$$H_2C{=}CH_2(g) + H_2O(g) \underset{\text{endothermic}}{\overset{\text{exothermic}}{\rightleftharpoons}} CH_3CH_2OH(g)$$

1. *If you increase the pressure, conditions will favour the forward reaction and ethanol (CH_3CH_2OH) will be formed. This is because two molecules of $H_2C{=}CH_2/H_2O$ react to form only one molecule of CH_3CH_2OH. This reduces the pressure.*
2. *Raising the temperature favours the reverse reaction. This is because it's endothermic (see page 43) and absorbs the extra heat energy, lowering the temperature.*

These observations can be summarised by an important rule known as Le Chatelier's Principle:

"A reversible reaction will move its equilibrium position to resist any change in the conditions."

(Remember that altering the pressure of a gas is equivalent to altering the concentration of a solution.)

Have a go at this question:

1) You are making ethanol from ethene and steam using the reaction shown above. What will happen to the yield of ethanol if you increase the amount of steam in the reaction mixture?

Answers

1) Increasing the amount of steam will increase the concentration of particles on the left of the equation (which will also increase the pressure on the left hand side), and move the position of equilibrium to the right, increasing the yield of ethanol.

Influence of Conditions on Yield

Deciding on the Best Conditions to Use

Thanks to Le Chatelier's principle (see page 27) you might think it would be easy to work out the optimum conditions for a reversible reaction. But in real life it's not quite that simple.

For most reversible reactions that are used on an industrial scale there are other factors, such as cost and time, that need to be taken into account.

Have a look at the conditions used for the production of ethanol from ethene:

$$H_2C{=}CH_2(g) + H_2O(g) \underset{\text{endothermic}}{\overset{\text{exothermic}}{\rightleftharpoons}} CH_3CH_2OH(g)$$

Catalyst: H_3PO_4
Temperature: 300 °C
Pressure: 60 atm

Temperature:

1) *Lowering the temperature would favour the forward reaction (see page 27), and so it should increase the yield of ethanol.*
2) *But lowering the temperature also means that fewer of the particles in the reaction mixture will have enough energy to react.*
   So lowering the temperature will slow down both the forward and reverse reactions.
3) *A low temperature would make the forward reaction too slow to be useful.*
   So a compromise temperature of 300 °C is used.

Pressure:

1) *Increasing the pressure would favour the forward reaction (see page 27), and increase the rate of reaction (as it increases the probability of collisions between particles). This would increase the yield of ethanol.*
2) *But producing high pressures uses a lot of energy and costs a lot of money.*
   You'd need some pretty strong equipment to stand up to the high pressures too — and that would be expensive to buy and maintain.
3) *To make the reaction economic, a moderately high compromise pressure of 60 atm is used.*

Have a go at these questions on the conditions used in the production of ethanol from ethene:

1) *Explain why it is not a good idea to run the reaction industrially at a temperature of 40 °C.*
2) *Explain why it is not a good idea to run the reaction industrially at a pressure of 500 atm.*

Answers

1) *The temperature is low, which would favour the forward reaction, and increase the yield of ethanol. But it is so low that the forward reaction rate will be much too slow to be economic.*
2) *The pressure is high, which would favour the forward reaction, and increase the yield of ethanol. But such a high pressure would be very expensive to maintain, making the reaction uneconomic.*

Formulae of Compounds

Deducing the Formulae of Ionic Compounds

The formula of a compound tells you the ratio of the elements that it contains. This ratio is fixed, and for ionic compounds that means it's easy to work out the formula from the charges on the ions.

Metal ions (and hydrogen ions) always carry a positive charge, whilst non-metal ions carry a negative charge. If you imagine that a positive charge is a 'hook' and a negative charge is an 'eye' then the formula can be deduced by exactly matching up the hooks and eyes. (This is to make the compound electrically neutral — it's the same idea as the ionic lattices on page 10.)

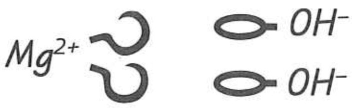
Na^+ (sodium ion) has +1 charge so 1 hook
Mg^{2+} (magnesium ion) has +2 charge so 2 hooks
OH^- (hydroxide ion) has -1 charge so 1 eye
O^{2-} (oxide ion) has -2 charge so 2 eyes

Example 1: What is the formula of sodium oxide?

We need an extra Na^+ to give us a second hook to match the second of the eyes on the O^{2-} ion.

We have 2 Na^+ ions to every O^{2-} ion, so the formula is Na_2O.

Example 2: What is the formula of magnesium hydroxide?



Note the use of a bracket to show 2 lots of OH which is not the same as OH_2. Brackets are most often used when the non-metallic ion contains more than one element.

There are 2 OH^- ions to every Mg^{2+} ion so the formula is $Mg(OH)_2$.

Now try these:

Use the charges on the ions at the bottom of the box to deduce the formulae of the following ionic compounds.

1) sodium chloride
2) calcium bromide
3) sodium carbonate
4) aluminium oxide
5) iron(II) chloride
6) potassium oxide
7) aluminium chloride
8) potassium nitrate
9) aluminium sulfate
10) iron(III) nitrate

aluminium: Al^{3+}	bromide: Br^-	calcium: Ca^{2+}	carbonate: CO_3^{2-}
chloride: Cl^-	iron(II): Fe^{2+}	iron(III): Fe^{3+}	nitrate: NO_3^-
oxide: O^{2-}	potassium: K^+	sodium: Na^+	sulfate: SO_4^{2-}

Answers

1) $NaCl$ 6) K_2O
2) $CaBr_2$ 7) $AlCl_3$
3) Na_2CO_3 8) KNO_3
4) Al_2O_3 9) $Al_2(SO_4)_3$
5) $FeCl_2$ 10) $Fe(NO_3)_3$

Writing and Balancing Equations

Rules for Working out the Products Formed in Reactions

To write a balanced equation you need to be aware of the different ways that compounds react. Many examples involve the reactions of acids to form salts. It'll help you to know some of the rules for working out what products are formed during particular reactions.

Making Salts:

1. If sulfuric acid is used the salt will be '________' sulfate where '________' is a metal.
   For example: magnesium + sulfuric acid = magnesium sulfate + hydrogen
2. If hydrochloric acid is used the salt will be '________' chloride where '________' is a metal.
   For example: magnesium + hydrochloric acid = magnesium chloride + hydrogen
3. If nitric acid is used the salt will be '________' nitrate where '________' is a metal.
   For example: magnesium + nitric acid = magnesium nitrate + hydrogen

(Sulfuric acid = H_2SO_4 Hydrochloric acid = HCl Nitric acid = HNO_3)

Some Other General Rules:

1. Combustion reactions result in the formation of oxides.
   For example: $2H_2 + O_2 \rightarrow 2H_2O$
2. When fuels burn, carbon dioxide (CO_2) and water (H_2O) are normally produced.
   For example: $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$
3. Atoms in gases often go round in pairs: H_2, N_2, O_2, Cl_2

Writing Balanced Equations

To write a balanced symbol equation where you're given the reactants there are 5 simple steps:

1. Write out the word equation first.
2. Write the correct formula for each compound below its name (see page 29).
3. Go through each element in turn, making sure the number of atoms on each side of the equation balances.
4. If you changed any numbers, do step 3 again.
5. Keep doing this until all the elements balance.

Doing the third step:

If the atoms in the equation don't balance you can't change the molecular formulae — only the numbers in front of them.

For example:

$$CaO + 2HCl \rightarrow CaCl_2 + H_2O$$

There are two Cl on the right of the equation, so we need to have two HCl on the left-hand side. This also doubles the number of hydrogen atoms on the left-hand side, so that the hydrogens balance as well. This always works. If you can't get an equation to balance then it's wrong.

The example on the next page uses these rules to write a balanced symbol equation. Read through it carefully and make sure you understand the process.

Writing and Balancing Equations

Example: Write a balanced equation for the reaction of magnesium with hydrochloric acid.

Step 1: Magnesium + hydrochloric acid → magnesium chloride + hydrogen

Step 2: $Mg + HCl \rightarrow MgCl_2 + H_2$

Step 3: $Mg + 2HCl \rightarrow MgCl_2 + H_2$

(the Mgs already balance,
put a 2 in front of HCl to balance the Hs and Cls. Check that all still balances.)

Now try this question:

1) Write a balanced symbol equation for the combustion of methane (CH_4) in oxygen. Step 1 has been done for you.

   Step 1: Methane + oxygen → carbon dioxide + water

Use everything that you've learnt on the last three pages to answer these questions:

2) Balance the symbol equations for the following reactions:
   a) $K + H_2SO_4 \rightarrow K_2SO_4 + H_2$
   b) $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
   c) $Na_2O + HCl \rightarrow NaCl + H_2O$
   d) $KOH + H_2SO_4 \rightarrow K_2SO_4 + H_2O$
3) Write balanced symbol equations for the following reactions.
   a) the complete combustion of the fuel ethanol (C_2H_5OH) in oxygen.
   b) the reaction of calcium hydroxide with hydrochloric acid to give calcium chloride and water.

chloride ion: Cl^- hydrogen ion: H^+ hydroxide ion: OH^- calcium ion: Ca^{2+}

Answers

1) Step 1: Methane + oxygen → carbon dioxide + water
   Step 2: $CH_4 + O_2 \rightarrow CO_2 + H_2O$
   Step 3: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
   (The Cs already balance, so put a 2 in front of H_2O to balance the Hs. Now put a 2 in front of O_2 to balance the Os. Check that all still balances.)
2) a) $2K + H_2SO_4 \rightarrow K_2SO_4 + H_2$
   b) $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
   c) $Na_2O + 2HCl \rightarrow 2NaCl + H_2O$
   d) $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$
3) a) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
   b) $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

Determination of Formulae from Reacting Masses

On page 29 you saw how to work out the formula of an ionic compound from the position of the reactants in the Periodic Table. Formulae are not always so easy to work out. Often, the only way to find out the formula of a compound is through experimentation and calculation.

The following three pages give explanations, examples and practice of the techniques used to find the formulae of various compounds. It will really help you to be comfortable with these methods, and able to do the calculations quickly.

Calculating Empirical Formulae

You can calculate the formula of a compound from the masses of the reactants.
Here is a simple set of rules to follow when calculating a formula:
1. Write the mass or percentage mass of each element
2. Divide by the relative atomic mass (A_r) of the atom
3. Divide all answers by the smallest answer
4. If required: multiply to make up to whole numbers
5. Use the ratio of atoms to write the formula (this gives the empirical formula — see p.33)

Example: Find the formula of an oxide of aluminium formed from 9g aluminium and 8g oxygen.

Using the rules above:	Al	O
1. Write the mass of each element:	9g	8g
2. Divide by the relative atomic mass of each atom:	9/27 = 0.333	8/16 = 0.5
3. Divide by the smallest answer:	0.333/0.333 = 1	0.5/0.333 = 1.50
4. Make up to whole numbers (×2):	1×2 = 2	1.5×2 = 3
5. Use the ratio of atoms to write formula:	Al_2O_3	

Now try these questions:

1) Find the empirical formulae of the following three oxides of lead:
   - the first contains 12.94g of lead to every 1g of oxygen
   - the second comprises 6.47g of lead to every 1g of oxygen
   - the third contains 9.7g of lead to every 1g of oxygen
   (Relative atomic mass values: Pb = 207, O = 16)
2) Calculate the empirical formula of the carboxylic acid that is comprised of 4.3% hydrogen, 26.1% carbon and 69.6% oxygen.
   (Relative atomic mass values: H = 1, C = 12, O = 16)

Answers

1) PbO
   PbO_2
   Pb_3O_4
2) CH_2O_2

Calculation of Empirical Formulae

Empirical and Molecular Formulae

The empirical formula of a compound is the simplest ratio of the atoms of each element in the compound.
The molecular formula of a compound gives the actual number of atoms of each element in the compound.
For example a compound with the molecular formula C_2H_6 has the empirical formula CH_3.
The ratio of the atoms is one C to every three Hs.

To find the molecular formula from the empirical formula, you need to know the relative formula mass (see page 3) of the compound. This will usually be given to you in the question. Read through the example below and then try the questions.

Calculating the molecular formula from the empirical formula:

Example: Calculate the molecular formula of a hydrocarbon molecule if the compound contains 85.7% carbon and it's relative formula mass is 42.

1. Calculate the empirical formula:

	C	H
	85.7	14.3
	85.7/12	14.3/1
	= 7.14	= 14.3
	7.14/7.14	14.3/7.14
	= 1	= 2
	CH_2	

2. Calculate how many multiples of the empirical formula the molecular formula contains:

1 empirical formula (CH_2) has a relative mass of 12 + 1 + 1 = 14
The molecular formula has a relative mass of 42.
42/14 = 3
So the molecular formula is 3 × the empirical formula
$C \times 3 = 3C$, $H_2 \times 3 = 6H$

The molecular formula is: C_3H_6

Have a go at these:

1) Calculate the molecular formula of a compound containing 52.2% carbon, 13.0% hydrogen and 34.8% oxygen if the relative formula mass of the compound is 46.
(Relative atomic mass values: C = 12, H = 1, O = 16)

2) Calculate the molecular formula of a hydrocarbon with a relative formula mass of 78 that contains 92.3% carbon.
(Relative atomic mass values: C = 12, H = 1)

Answers
1) C_2H_6O
2) C_6H_6

Determination of Formulae from Reacting Masses

More Practice Calculations

The more practice you get at these calculations the better. Be careful — some questions are asking you for the empirical formula, and some for the molecular formula.

Have a go at these:

1) *What is the ratio of carbon to hydrogen atoms in a hydrocarbon molecule containing 75% carbon by mass?*
   (Relative atomic mass values: C = 12, H = 1)
2) *What is the empirical formula of an oxide of nitrogen comprised of 30.4% nitrogen by mass?*
   (Relative atomic mass values: N = 14, O = 16)
3) *What is the empirical formula of an oxide of sulfur if a 12g sample contains 6g of sulfur?*
   (Relative atomic mass values: S = 32, O = 16)
4) *A compound of oxygen, titanium and iron contains 28% iron and 36% titanium. What is the simplest formula of the compound?*
   (Relative atomic mass values: Fe = 55.8, Ti = 47.9, O = 16.0)
5) *Calculate the molecular formula of an oxide of phosphorus that contains 0.775g of phosphorus to every 1g of oxygen and which has a relative formula mass of 284.*
   (Relative atomic mass values: P = 31, O = 16)
6) *Calculate the molecular formula of aluminium chloride if it has 3.94g of chlorine to every 1g of aluminium, and a relative formula mass of 267.*
   (Relative atomic mass values: Al = 27, Cl = 35.5)
7) *Calculate the molecular formula of titanium oxide if it contains one third oxygen by mass and has a relative formula mass of 144.*
   (Relative atomic mass values: Ti = 48, O = 16)

Answers

1) 1:4 (CH_4)
2) NO_2
3) SO_2
4) $Fe_2Ti_3O_9$
5) P_4O_{10}
6) Al_2Cl_6
7) Ti_2O_3

Making Use of the Periodic Table

The Periodic Table Holds Lots of Useful Information

From the <u>atomic number</u> you can work out the <u>electronic configuration</u> of an atom of the element.

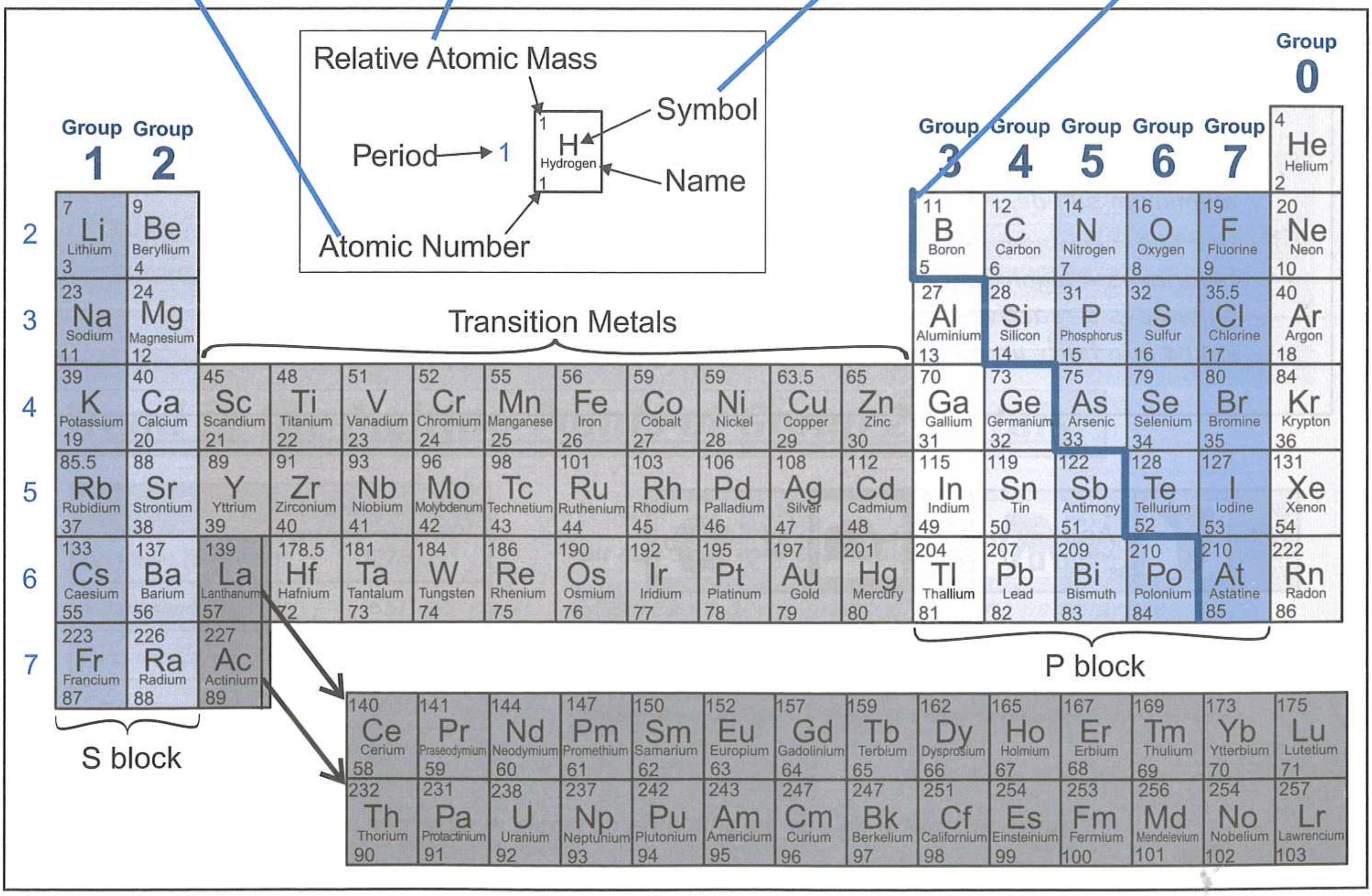
The <u>mass number</u> is the total number of protons and neutrons in the nucleus of one atom of an element.

<u>Atomic number</u> (or proton number) is the number of protons in the nucleus. This is also the number of electrons in an atom.

The <u>relative atomic mass</u> is the <u>average mass number</u> for all the <u>isotopes</u> of an element, taking into account their relative abundance (see page 3).

The <u>first letter</u> of the symbol is always a capital. If there is a second then it is lower case e.g. C, Na, K, Cl

This line marks the <u>boundary</u> between <u>metals and non-metals</u>. All those elements to the right of it are non-metals.



You can follow trends in physical and chemical properties down <u>VERTICAL GROUPS</u> and across <u>HORIZONTAL PERIODS</u>.

All elements in a group have the same <u>outer electron configuration</u>, and so form ions with the same charge. Knowing the number of electrons in the outer shell means you can work out the <u>formulae of compounds</u>.

Making Use of the Periodic Table

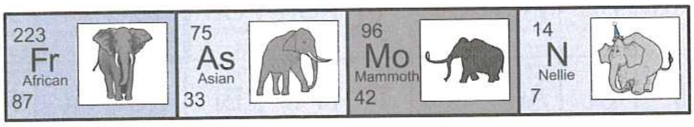
Learn and Practise Some Important Skills

You'll find chemistry heaps easier if you get used to working with the basic information that's contained in the Periodic Table.

1) *One of the most important things that you can learn is the names and symbols of the elements — particularly the more common ones that you'll need to use a lot.*
2) *If you know the atomic number of an element in the Periodic Table, you can work out its electronic configuration.*
3) *Being able to do that will help you to work out the formulae of compounds, ionic or covalent, quickly.*

Use these questions and the Periodic Table on the previous page to improve your skills.

1) *Find the symbols or names of the following elements: calcium, vanadium, phosphorus, Br, tin, Au, W, potassium, manganese, boron, Sb, thallium.*
2) *Find the proton number for each of the elements in question 1.*
3) *Using only the proton number, write out the electronic configurations, using both crosses and the shorthand (2,8,4), for these elements: Na, S, Ca, N, Mg, He.*
4) *What is the charge on the ions formed by each of these elements: K, magnesium, nitrogen, sulfur, Al, I?*
5) *Work out the formulae of the following compounds: magnesium oxide, lithium bromide, aluminium sulfide, iron(II) oxide, copper(II) chloride.*
6) *Use the following information to predict the properties of bromine.*
   Fluorine is a highly reactive gas with a boiling point of –188 °C,
   chlorine is a reactive gas with a boiling point of –35 °C,
   iodine is a fairly unreactive solid with a boiling point of 184 °C.



Answers

1) Ca, V, P, bromine, Sn, gold, tungsten, K, Mn, B, antimony, Tl.
2) In order: 20, 23, 15, 35, 50, 79, 74, 19, 25, 5, 51, 81.
3) Shorthand electronic configurations: Na 2,8,1; S 2,8,6; Ca 2,8,8,2; N 2,5; Mg 2,8,2; He 2. Use these to check your drawn out electronic configurations.
4) In order: +1, +2, –3, –2, +3, –1.
5) In order: MgO, LiBr, Al_2S_3, FeO, $CuCl_2$.
6) The actual properties of bromine are, it is a fairly reactive liquid, which has a boiling point of 59 °C. If you were close then well done.

Reactivity and Group 2

Trend in Reactivity Down the Group

During their reactions, Group 2 metals donate their two outer electrons to another atom (usually a non-metal). The reactivity of Group 2 metals depends on how easily the outer electrons can be donated. The easier the electrons can be donated, the more reactive the metal will be.
By experimenting, you can find that:

Reactivity increases as you go down Group 2.

To see why this is, think about the factors that affect how strongly an electron is held by the nucleus:

1) The first is the positive nuclear charge. The nuclear charge attracts electrons and keeps them in orbit. A greater nuclear charge provides a stronger force of attraction and makes it more difficult for the atom to donate its outer electrons. As you go down the group, the nuclear charge increases, so if this was the only factor, reactivity would decrease down Group 2. But that isn't the case.
2) The second factor is that in larger atoms, the outer electrons are further away from the positive nucleus. The electrostatic attraction quickly decreases in strength with distance from the source.
3) The third factor is electron shielding. As the atoms in Group 2 get larger, the number of full electron shells round the nucleus increases. These negative charges shield the two outer electrons from the attraction of the positive nucleus.

The increase in the distance between the outer electrons and the nucleus, and the increased shielding as you go down the Group, far outweigh the increase in nuclear charge.

Trend in the Melting Points of Group 2 Metals

	Melting Point (°C)
Beryllium (Be)	1278
Magnesium (Mg)	651
Calcium (Ca)	839
Strontium (Sr)	769
Barium (Ba)	727

You can see from the table that:

As you go down Group 2, melting point decreases.

Magnesium doesn't fit in with the general trend. It behaves a bit oddly because it has a slightly different structure to the other Group 2 metals.

This is also due to the increase in electron shielding as you go down the group.
Group 2 metals, like all other metals, are held together in a lattice structure by metallic bonds. Metallic bonds are formed when the outer electrons from each atom break free from the nucleus, leaving positive ions and free electrons.
The strength of the metallic bonds depend on how strong the attraction is between the positive ions and the free electrons. The more shielded the positive nuclei are, the weaker the attraction will be, and so the less energy will be required to break the bond and melt the metal.

Reactivity and Group 7

Some General Properties of Group 7 Elements

Group 7 elements all have 7 electrons in their outer shell.
As a result these elements either:

1) Form ionic compounds by gaining an extra electron

or 2) Share a pair of electrons and form a covalent bond

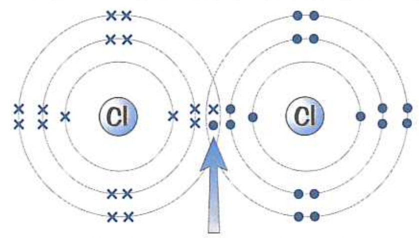
In their elemental state, the halogens bond covalently, forming diatomic molecules (two atoms joined with a single covalent bond) with coloured vapours:

Chlorine (Cl_2) is a yellow/green gas at room temperature

Bromine (Br_2) is a brown liquid at room temperature

Iodine (I_2) is a grey solid at room temperature (which sublimes to produce a purple vapour)

In each case the atoms share an electron pair, as shown below:



single covalent bond

As you go down the group:
1) melting point increases
2) boiling point increases

This is because the strength of the weak attraction between molecules increases as the number of electrons in the molecules increase.

Trend in Reactivity Down Group 7

During their reactions, Group 7 elements attract an extra electron from another atom.
The reactivity of Group 7 elements depends on how strongly the nucleus can attract electrons.
The stronger the attraction, the more reactive the element will be.

Reactivity decreases as you go down Group 7.

The same arguments apply to the reactivity trend in the Group 7 elements as the one in Group 2.

1) As with the Group 2 elements, nuclear charge increases as you go down the group.
A greater nuclear charge will attract the extra electron required to fill the outer shell more strongly. This works to increase the reactivity of the elements as you go down the group.
2) However, as the atoms get bigger, the extra shells of electrons shield the nuclear charge more effectively. So the nucleus is less able to attract the extra electron the atom wants.

In Group 7 this shielding outweighs the effect of increasing nuclear charge. The elements at the top of the group are best able to attract an extra electron, and are more reactive.

Reactivity Trends in Groups 2 and 7

Group 7 Reactivity and Displacement Reactions

You can demonstrate the relative reactivity of the Group 7 elements using displacement reactions.

If you mix a halogen with a solution containing halide ions, a more reactive halogen will displace a less reactive halide ion (one below it in the group) from solution.

e.g. Fluorine is more reactive than chlorine.

$$F_2\ (aq) + 2Cl^-\ (aq) \rightarrow Cl_2\ (aq) + 2F^-\ (aq)$$

The chloride ions have been displaced from the solution.

Use the trends from the previous pages to answer the following questions:

1) The following are descriptions of the reactions of Be and Ca with cold water.
   Use them to predict the reactions of Mg and Sr.
   - Beryllium will not react with cold water at all
   - Calcium reacts steadily with cold water to produce hydrogen gas and calcium hydroxide.
2) Explain what would happen if you mixed the following halogens and halide solutions.
   a) chlorine and bromide
   b) bromine and iodide
   c) iodine and chloride
   d) iodine and bromide
   e) chlorine and iodide
3) For each of the reactions that takes place in question 2, write out an ionic equation.

Answers

1) Magnesium reacts very slowly with cold water to produce hydrogen gas and magnesium hydroxide.
   Strontium reacts vigorously with cold water to produce hydrogen gas and strontium hydroxide
2) a) Chlorine is more reactive than bromine, so it displaces the bromide in solution producing chloride and bromine.
   b) Bromine is more reactive than iodine, so it displaces the iodide in solution producing bromide and iodine.
   c) Iodine is less reactive than chlorine, so no reaction takes place.
   d) Iodine is less reactive than bromine, so no reaction takes place.
   e) Chlorine is more reactive than iodine, so it displaces the iodide in solution producing chloride and iodine.
3) $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$
   $Br_2 + 2I^- \rightarrow 2Br^- + I_2$
   $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$

Reaction Types

There are lots of types of chemical reaction. You will need to know all of them quite well. More importantly you need to be able to recognise when an individual reaction can be classed as more than one type. These pages give you types, explanations and examples (in alphabetical order).

Addition

This is a reaction in which atoms are added to an unsaturated bond so that the bond becomes saturated.

e.g. ethene (C_2H_4) + H_2O → ethanol (C_2H_5OH)

Combustion

This is the chemical reaction between a fuel and oxygen. Normally the fuel is an organic compound and the products are carbon dioxide and water. Without enough oxygen, incomplete combustion takes place producing poisonous carbon monoxide (see page 20).

Condensation

This is similar to an addition reaction in which a simple molecule like water is also formed.

Cracking

This is the (thermal) decomposition of long-chain hydrocarbon molecules from crude oil into shorter- chain alkanes and alkenes. This requires high temperatures and pressures and a catalyst (usually aluminium oxide), and makes hydrocarbons that are more useful.

e.g. decane ($C_{10}H_{22}$) → octane (C_8H_{18}) + ethene (C_2H_4)

Dehydration

This is the removal of water from a compound by heating.
In organic molecules it usually results in the formation of a C=C bond.

e.g. ethanol (C_2H_5OH) → ethene (C_2H_4) + H_2O

Displacement

This is a reaction where one element displaces another, less reactive, element from a compound. This usually takes place between metals, but also with halogens.

e.g. $2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$ (the Thermite Reaction)

Disproportionation

This is a rare type of chemical reaction where an element in a reactant is oxidised and reduced at the same time. Chlorine can undergo disproportionation reactions.

e.g. $Cl_2(aq) + H_2O(l) \rightarrow HOCl(aq) + HCl(aq)$

HOCl(aq): Chloric(I) acid; HCl(aq): Hydrochloric acid

The chlorine has been: oxidised reduced

Reaction Types

Electrolysis

This is a process that uses electricity to break down a compound. The reactant or reactants must be in the liquid state — either molten or in solution. The particles have to be able to move. The name can be split into two bits: 'electro-' for electricity and '-lysis' means to break down. An example is the electrolysis of bauxite to obtain pure aluminium.

Elimination

This is just the removal of a small molecule from a larger molecule. Usually H_2O or H_2 is removed (and not replaced by anything else).

Endothermic

Any chemical reaction that takes in heat energy. This means that the reactants will have less energy than the products.

Exothermic

Any chemical reaction that gives out heat energy. This happens because the products have less energy than the reactants. Hint: 'exo-' and 'exit' come from the same word meaning 'out'.

Hydrogenation

This is the addition of a molecule of hydrogen (H_2) across a C=C bond. One atom attaches to each carbon.

e.g. ethene (C_2H_4) + H_2 → ethane (C_2H_6)

Neutralisation

This is the reaction between a basic compound and an acid. The products always include the salt of the acid, water and other products dependent on the acid and base.

e.g. $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

Oxidation

There are two possible definitions for this; the best is the loss of electrons. Another useful one is the gain of oxygen. It is the opposite of reduction.

Precipitation

A precipitate is a solid that is formed in a solution by a chemical reaction or by a change in temperature affecting solubility. Precipitates are insoluble in the solvent. A precipitation reaction is simply any reaction that produces a precipitate.

Reaction Types

REDOX

This is the name for a reaction that involves both reduction and oxidation processes. It is usually used to describe reactions that just involve electron transfer.

e.g. $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

(reduction: $Cu^{2+}(aq) \rightarrow Cu(s)$; oxidation: $Fe(s) \rightarrow Fe^{2+}(aq)$)

Reduction

There are two possible definitions for this; the best is the gain of electrons. The other useful one is the loss of oxygen. Important point: oxidation and reduction ALWAYS happen together — it is impossible to have one without the other.

Reversible

This is the name given to any chemical reaction that can go forwards and backwards at the same time. That means that the reactants will form the products, but that the products will also react (or decompose) to give the reactants.

e.g. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Substitution

This is simply a reaction in which an atom (or group of atoms) in a molecule are swapped for a different atom (or group of atoms).

Thermal Decomposition

This is where one compound breaks down, under heating, into two or more simpler compounds. A classic example is the breakdown of any carbonate compound,

e.g. $CaCO_3 \xrightarrow{heat} CaO + CO_2$

Cracking of hydrocarbons is also an example.

Now try these questions:

Write down all the different types of reaction that each of the following could be classed as.

1) burning ethanol
2) iron + copper sulfate → iron sulfate + copper
3) hydrochloric acid + sodium hydroxide → sodium chloride + water (gets hotter)
4) propene (C_3H_6) + H_2 → propane (C_3H_8)

Answers

1) combustion, exothermic, oxidation, REDOX, (reduction)
2) displacement, oxidation, precipitation, REDOX, reduction, substitution
3) exothermic, neutralisation, oxidation, substitution, REDOX, reduction
4) addition, hydrogenation, (oxidation, REDOX, reduction)

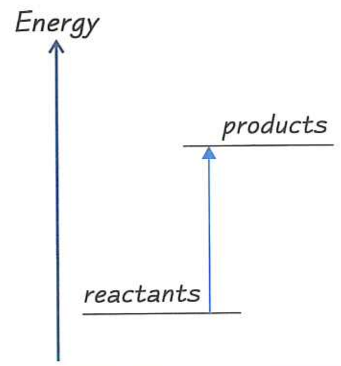
Exothermic and Endothermic Reactions

In an <u>exothermic</u> reaction, heat energy is given out (the room temperature rises).
In an <u>endothermic</u> reaction, heat energy is taken from the surroundings (the room temperature drops).

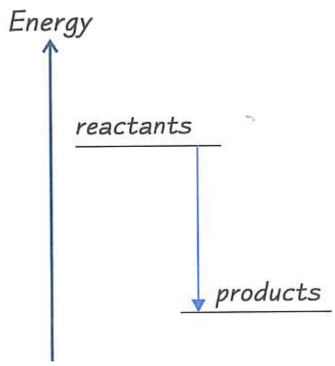
Making and Breaking Bonds

When two atoms joined by a covalent bond are <u>separated</u>, the energy required to do this must be provided from the surroundings. Reactions which require an <u>energy input</u> like this are called <u>endothermic</u> reactions.
When two atoms become <u>joined</u> together by forming a covalent bond, energy is released to the surroundings. Reactions that <u>release energy</u> to the surroundings are called <u>exothermic</u> reactions.

Reactions can be Represented by Energy Level Diagrams



In an <u>endothermic</u> reaction, the reactants absorb energy from the surroundings. The products therefore have more energy than the reactants.



In an <u>exothermic</u> reaction, the reactants release energy to the surroundings. The products therefore have less energy than the reactants.

Have a go at these questions:

Are the following reactions exothermic or endothermic?

1) photosynthesis (uses energy from sunlight)
2) respiration
3) burning coal
4) sodium hydrogencarbonate + hydrochloric acid (temperature drops)
5) acid + hydroxide (gets hotter)
6) methane + steam (cools as they react)

<u>Answers</u>
1) Endothermic
2) Exothermic
3) Exothermic
4) Endothermic
5) Exothermic
6) Endothermic

Calculations Involving Making and Breaking Bonds

Average Bond Energy

Bonds between different atoms require different amounts of energy to break them. When the same two atoms bond in the same way, the amount of energy needed is always about the same. The average bond energy values for some common bonds are given below:

C — H	413	C — O	360	C = C	612
O = O	498	H — H	436	C = O	743
C — C	348	O — H	463		

All values are in $kJmol^{-1}$ (that means the same as kJ/mol).

The values tell you that:

e.g. It takes 413 kJ of energy to break 1 mole of C — H bonds.
It takes 463 × 2 = 926 kJ to break 1 mole of water (which has 2 O — H bonds per molecule) into oxygen and hydrogen atoms.
743 × 2 = 1486 kJ are released when 1 mole of CO_2 (which has 2 C = O bonds) forms.

Calculating the Change in Energy

When a reaction takes place, the change in energy is simply:

sum of energy required to break old bonds – sum of energy released by new bonds formed

Example: Calculate the energy change involved when 1 mole of methane burns in oxygen.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

(The equation tells you that 1 mole of methane reacts with 2 moles of oxygen to form 1 mole of carbon dioxide and 2 moles of water.)

Solution:

Step 1: Calculate the energy required to break all of the bonds between the reactant atoms

4 C — H bonds	=	4 × 413	= 1652 kJ
2 O = O bonds	=	2 × 498	= 996 kJ
		TOTAL	= 2648 kJ

Remember — there are 2 moles of oxygen

Step 2: Calculate the energy released by all the new bonds formed between the atoms in the products

2 C = O bonds	=	2 × 743	= 1486 kJ
4 O — H bonds	=	4 × 463	= 1852 kJ
		TOTAL	= 3338 kJ

1 mole of CO_2 is formed and 2 moles of H_2O

Step 3: Combine the two values to give an overall value for the energy change

The overall energy change combines +2648 and –3338 which equals –690 $kJmol^{-1}$

The negative sign shows that energy is being released to the surroundings, indicating that this is an exothermic reaction. This is what you would expect, since this is a combustion reaction.

Calculations Involving Making and Breaking Bonds

Calculate how much energy is released in the following reactions:

1) burning 1 mole of propane $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
2) burning 1 mole of ethanol $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
3) burning 1 mole of octane $C_8H_{18} + 12\frac{1}{2}O_2 \rightarrow 8CO_2 + 9H_2O$
4) hydrogenation of 1 mole of ethene $C_2H_4 + H_2 \rightarrow C_2H_6$

Answers

1) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms

8 C — H bonds = 8 × 413 = 3304
2 C — C bonds = 2 × 348 = 696
5 O = O bonds = 5 × 498 = 2490
TOTAL = 6490

Step 2: Calculate the energy released by all the new bonds formed between the product atoms

6 C = O bonds = 6 × 743 = 4458
8 O — H bonds = 8 × 463 = 3704
TOTAL = 8162

Step 3: Combine the two values to give an overall value for the energy change

+6490 − 8162 = −1672 $kJmol^{-1}$

2) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms

1 C — C bond = 1 × 348 = 348
1 C — O bond = 1 × 360 = 360
5 C — H bonds = 5 × 413 = 2065
1 O — H bonds = 1 × 463 = 463
3 O = O bond = 3 × 498 = 1494
TOTAL = 4730

Step 2: Calculate the energy released by all the new bonds formed between the product atoms

4 C = O bonds = 4 × 743 = 2972
6 O — H bonds = 6 × 463 = 2778
TOTAL = 5750

Step 3: Combine the two values to give an overall value for the energy change

+4730 − 5750 = −1020 $kJmol^{-1}$

3) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms

18 C — H bonds = 18 × 413 = 7434
7 C — C bonds = 7 × 348 = 2436
12.5 O = O bonds = 12.5 × 498 = 6225
TOTAL = 16 095

Step 2: Calculate the energy released by all the new bonds formed between the product atoms

16 C = O bonds = 16 × 743 = 11 888
18 O — H bonds = 18 × 463 = 8334
TOTAL = 20 222

Step 3: Combine the two values to give an overall value for the energy change

+16 095 − 20 222 = −4127 $kJmol^{-1}$

4) Step 1: Calculate the energy required to break the H — H and C = C bonds

1 H — H bond = 1 × 436 = 436
1 C = C bond = 1 × 612 = 612
TOTAL = 1048

Step 2: Calculate the energy released by all the new bonds formed between product atoms

2 C — H bonds = 2 × 413 = 826
1 C — C bond = 1 × 348 = 348
TOTAL = 1174

Step 3: Combine the two values to give an overall value for the energy change

+1048 − 1174 = −126 $kJmol^{-1}$

Evaluating Data

Evidence is Reliable if it Can be Repeated

Scientific evidence needs to be reliable (or reproducible). If it isn't, then it doesn't really help you.

When you're doing an investigation, you need to repeat your experiment several times to make sure your results are reliable — you should get round about the same answer each time.

RELIABLE means the results can be consistently reproduced in independent experiments.

Example: Cold Fusion.

In 1989, two scientists claimed that they'd produced 'cold fusion' (the energy source from the Sun at room temperature). It was huge news — if true, it could have meant clean energy from sea water. But other scientists just couldn't get the same results — they weren't reliable. And until they are, the scientific community won't take cold fusion seriously.

Repeating an Experiment Lets You Find a Mean Result

If you repeat an experiment, your results will usually be slightly different each time you do it. You can use the mean (or average) of the measurements to represent all these values. The more times you repeat the experiment the more reliable the average will be.
To find the mean:

ADD TOGETHER all the data values then DIVIDE by the total number of values in the sample.

Example: The rate of reaction between hydrochloric acid and magnesium metal.

You're doing the rate of reaction experiment shown on page 24. You repeat it three times to check your results are reliable. You get these results for the volume of gas given off after 30 seconds.

Run 1	Run 2	Run 3	Mean
23 cm^3	22 cm^3	25 cm^3	**23.3 cm^3**

⇐ $(23 + 22 + 25) \div 3 = 23.3\ cm^3$

Watch out for weird results that stick out like a hedgehog in a tea cup. These are called anomalous results. Think about what's likely to have caused them. For example — if one of the results above was only 5 cm^3, then something probably went wrong. Maybe the plunger got stuck. You should ignore the anomalous result when you calculate the mean.

Repeating experiments may not make your data more accurate. For instance, if your balance always reads 2 grams below the actual mass, repeating the measurements won't make it any more accurate.

Evidence Also Needs to be Valid

Collecting reliable data is important, but if the data doesn't answer your original question, it won't be any use. You need to think about what data to collect to make sure your results will be valid.

VALID means that the data is reliable AND answers the original question.

To answer scientific questions, scientists often try to link changes in one variable with another. For your data to be valid, you have to control all the other variables...

Evaluating Data

You Need to Control All the Variables

Scientists control the variables so that the only one that changes is the one they're investigating — all the others are kept constant. If one variable changes when another variable does, the variables are said to be correlated.

> Example: The effect of surface area on reaction rate.
>
> If you're investigating the effect of surface area on the rate of reaction you have to keep everything else, such as temperature and concentration, exactly the same. Surface area is the only variable that you change.

In experiments like this, you can say that one variable causes the other one to change because you have made sure that nothing else could be causing the change.

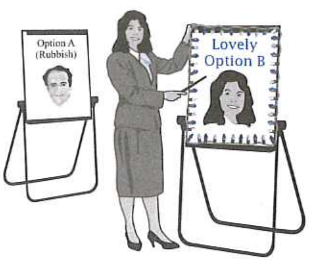
Controlling All the Variables Is Often Really Hard

The difficulty with a lot of scientific investigations is that it's very hard to control all the variables that might (just might) be having an effect.

> Example: Investigating whether chlorinated water increases cancer risk.
>
> Some studies claim that drinking chlorinated tap water increases the risk of certain cancers. But it's hard to control all the variables between people who drink tap water and people who don't. So designing a fair test is very tricky.

Even if some studies do show that people who drink more chlorinated water are slightly more likely to get certain cancers, it doesn't mean that drinking chlorinated water causes cancer. There will be heaps of differences between the groups of people. It could be due to any of them.

Correlation DOESN'T always mean cause



Watch Out For Bias

People who want to make a point sometimes present data in a biased way to suit their own purposes. They don't necessarily lie, they might just use the bits of data that support their argument, or phrase things in a leading way.

There are loads of reasons they might want to do this — for example, a company that sells water filters might tell you about the studies that found a link between drinking chlorinated water and cancer, and forget to mention all the studies that didn't find a link. They'll sell more water filters if people think drinking tap water isn't safe.

Benefits and Risks Must be Weighed Up

Scientific discoveries are often really useful, and they've improved our lives no end. But they often have risks attached. Society has to weigh things up and decide if they want to take the risks to get the benefits.

> Example: Adding chlorine to drinking water.
>
> Even if chlorine in water can cause health risks, these are extremely small compared to the risk of getting a nasty disease from drinking water with bacteria in.

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